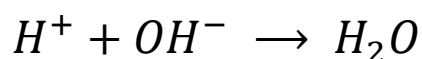
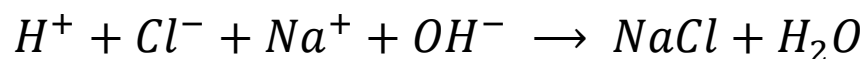


Practical Physical Chemistry  
Chemistry Department – College of Education  
Second Stage

**Experiment (1): Calculation of Heat of neutralization by calorimeter.**

In dilute solution strong acids and strong bases are completely dissociated to their ions. Furthermore, their salts are also completely dissociated in the solution. The neutralization of a strong acid by strong base may thus be written:



The heat effect is therefore independent of the nature of the anion of the acid and the cation of the base. The heat of neutralization may be determined by simple calorimeter.

## Apparatus and chemicals:

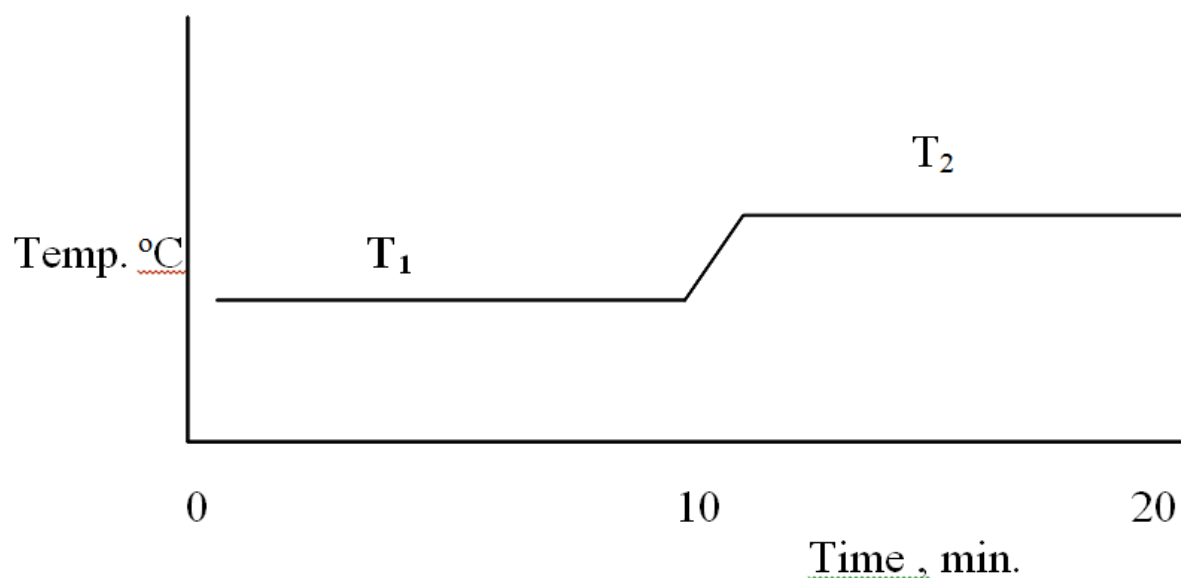
Calorimeter, thermometer, stirrer, 0.1N KOH, 2N HCl

## Procedure:

- 1- 100ml of 0.1N KOH is placed in a calorimeter, with continuous stirring record the temperature of the KOH solution every minute until 10min.
- 2- Add 16ml of 2N HCl to KOH solution and immediately record the temperature every minute until 10min.

## Calculation:

- 1- Plot a graph between temp. °C and time then find  $\Delta T$



2- Calculate (q)

$$q = -(m_1s_1 + m_2s_2)\Delta T$$

$m_1, m_2$  = mass of calorimeter (with thermometer and stirrer) and solution respectively.

$s_1, s_2$  = specific heat of calorimeter and solution respectively.

3- Find the amount (weight) of KOH present in 100ml of 0.1N, let this be  $x$

4- Calculate  $\Delta H$ .

$$\Delta H = \frac{q}{x} \times M.wt_{KOH}$$

## Experiment (2): Calculation of Heat of solution.

### Theory

The heat of solution is defined as : the change in enthalpy when one mole of a substance is dissolved in a specified quantity of solvent at a given temperature.

$$q = -(m_1s_1 + m_2s_2)\Delta T$$

### Apparatus and chemicals:

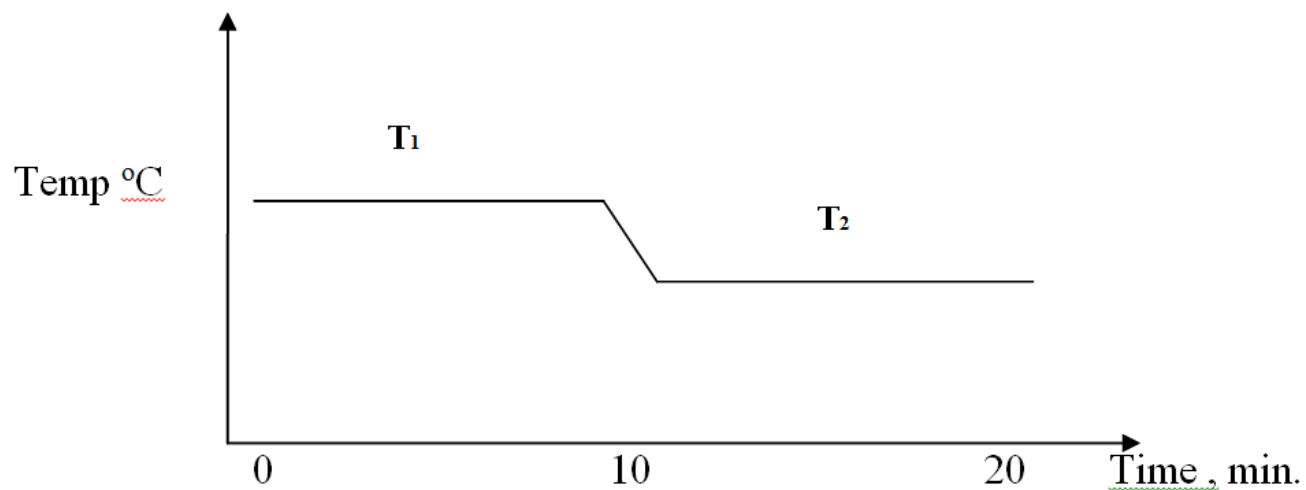
Calorimeter, thermometer, stirrer,  $KNO_3$ .

### Procedure:

- 1- 100ml of distilled water is placed in the calorimeter, with continuous stirring record the temperature of the water every minute until 10min.
- 2- Add 3.5g of  $KNO_3$  to the water and immediately record the temperature, then take it every minute until 10min.

**Calculation:**

1- Plot a graph between temp. °C and time then find  $\Delta T$



2- Calculate (q)

$$q = -(m_1s_1 + m_2s_2)\Delta T$$

3- Calculate  $\Delta H$ .

$$\Delta H = \frac{q}{x} \times M.wt_{KNO_3}$$

Where  $x$  is the weight of  $KNO_3$ .

### Experiment (3): Determination of Heat of solution from solubility.

#### Theory

The variation of the solubility of a substance with temperature is given by the relation:

$$\frac{d \ln S}{dT} = \frac{\Delta H_{sol.}}{RT^2} \dots \dots \dots (1)$$

Where  $S$  is the solubility of the substance and  $\Delta H_{sol.}$  is the heat of solution. This relationship may be regarded as a form of the van't Hoff isochore.

Assuming  $\Delta H_{sol.}$  is constant between the temperature  $T_1$  and  $T_2$  and integrating equation (1) between these limits:

$$\ln \frac{S_1}{S_2} = \frac{\Delta H}{R} \left( \frac{1}{T_2} - \frac{1}{T_1} \right) \dots \dots \dots (2)$$

Where  $S_1$  is the solubility at the absolute temperature  $T_1$ , and  $S_2$  is the solubility at the absolute temperature  $T_2$ .

If the solubility of a substance is determined at low different temperature; its heat of solution may be calculated by application of equation (2). If more accurate data are required solubility may be

determined at several different temperature and  $\Delta H$  may be obtained from a plot of  $\ln S$  against  $\frac{1}{T}$ .

### **Apparatus and chemicals**

Thermostats at 30, 40, 50, and 60 °C , benzoic acid potassium hydroxide (0.1N), pipette(25ml), phenolphthalein indicator, cotton wool .

### **Procedure**

- 1- 250ml of KOH (0.1N) must be prepared.
- 2- 1gm of benzoic acid is placed in a beaker which contains 100ml of hot water at about 70°C.
- 3- The beaker is placed in the thermostat at 60°C and leaving it for 15min. for thermal equilibrium.
- 4- A small piece of rubber tubing containing a cotton wool is attached to a 25ml pipette and 10ml of solution is withdrawn. (Note: the pipette may be warmed before use in order to prevent the crystallization of the solutions at the higher temperature)

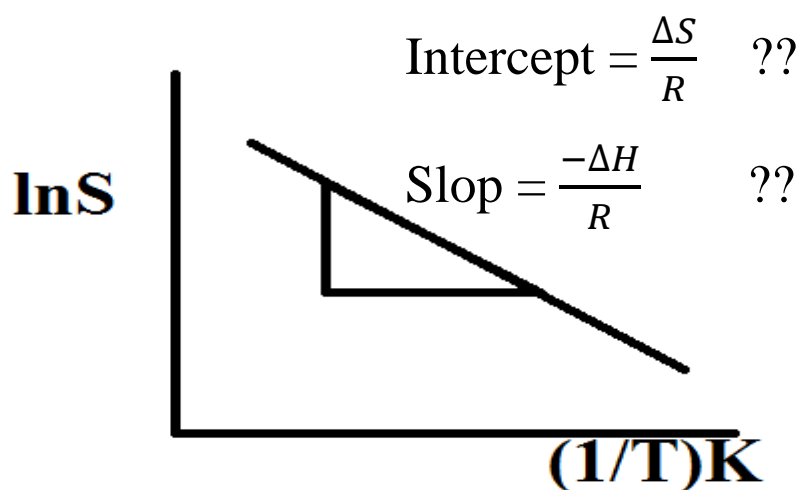
- 5- Put 10ml of solution in a conical flask which contain 30ml of hot water, then by using phenolphthalein as an indicator, titrate the solution with (0.1N) KOH and record the volume of KOH
- 6- Repeat these procedures from (3-5) by changing the temperature as (50,40 and 30°C).

## Calculation

- 1- Calculate the concentration of benzoic acid at each temperature (S)

$$(N.V)_{KOH} = (N.V)_{Benzoic\ Acid} \rightarrow N_{B.A} = S$$

- 2- Draw the plot between  $\ln S$  versus  $\frac{1}{T}$ .



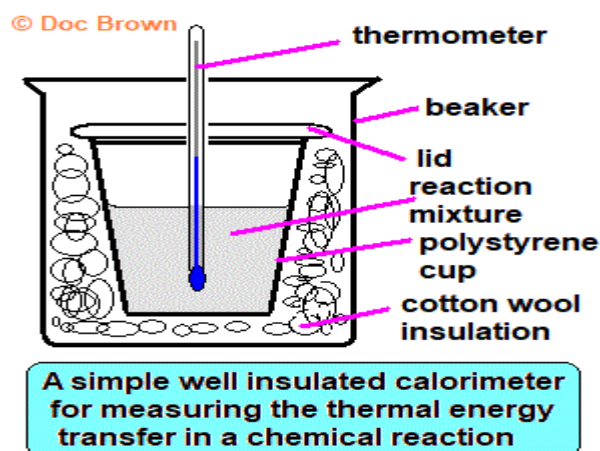


## Experiment (4): Determining the enthalpy of combustion of an alcohol.

Heat of combustion is the standard reaction enthalpy for the complete oxidation of an organic compound to  $CO_2$  gas and liquid  $H_2O$  if the compound contains C, H, and O.

Simple polystyrene calorimeter of low heat capacity can be used for any non-combustion reaction that will happen spontaneously at room temperature involving solutions or solid reacting/dissolving with/in a liquid like water and it doesn't matter if the reaction is exothermic or endothermic.

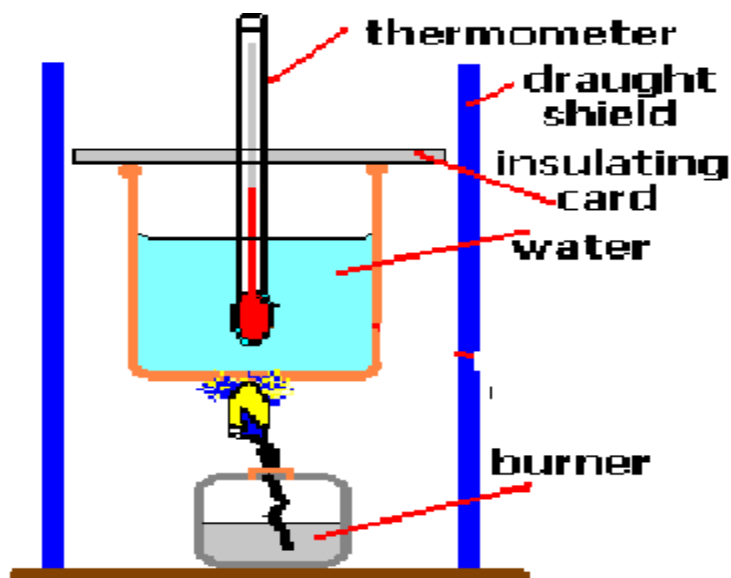
A double polystyrene cup system provides good thermal insulation for the system.



### Apparatus and chemicals:

Beaker or Conical flask, Thermometer, Spirit burner.

Arrange the equipment as shown in the following figure:



### Procedure:

- 3- Put 100 ml of water (100g) into the beaker (or conical flask).
- 4- The spirit burner contained the fuel ethanol  $\text{CH}_3\text{CH}_2\text{OH}$  and weighed 10g at the start.
- 5- Light the wick until the water temperature reaches  $\sim 70^\circ\text{C}$ .

### Calculation:

- 1- Calculate  $\Delta T$  of water (before and after heating).
- 2- Calculate  $\Delta m$  of ethanol (before and after burning).
- 3- Calculate the enthalpy of combustion of ethanol.

$$\text{Heat lost} = -\text{Heat gained}$$

$$Q = m \cdot S \cdot \Delta T$$

$$Q = \text{Thermal energy (J)}$$

$m = \text{mass of water}$

$S = \text{specific heat of water} \left( 1 \text{ cal, or } 4.18 \frac{\text{J}}{\text{g} \cdot ^\circ\text{C}} \right)$

$\Delta T = \text{temperature change of water}$

$$\Delta H_c \text{ of ethanol} = \frac{Q}{n}$$

$n = \text{mole of ethanol}$

**Note:**

**We can calculate the heat of combustion or the efficiency of some other fuels.**

<https://www.youtube.com/watch?v=DWNsZfyFLr4>

## **Experiment (5): Calculation of the heat capacity of different calorimeters.**

Specific heat is defined by the amount of heat needed to raise the temperature of 1 gram of a substance 1 degree Celsius ( $^{\circ}\text{C}$ ). Water has a high specific heat capacity which we'll refer to as simply "heat capacity", meaning it takes more energy to increase the temperature of water compared to other substances. This is why water is valuable to industries and in your car's radiator as a coolant.

The Zeroth Law of Thermodynamics states: "If two samples of matter, initially at different temperatures ( $T_{\text{H}}$  and  $T_{\text{C}}$ , respectively), are placed in thermal contact, heat will be lost by the hotter sample ( $T_{\text{H}}$ ) and gained by the cooler one ( $T_{\text{C}}$ ). This exchange of heat will take place until both samples achieve the same final temperature,  $T_{\text{F}}$ .

### **Apparatus and chemicals:**

Different calorimeter (Steel, Al, Cu, Glass)

### **Procedure:**

- 1- Mix 100 gm of Hot water (about  $65^{\circ}\text{C}$ ) with 100g ordinary water (at lab. Temperature) into cup calorimeter.
- 2- Measure the temperature of the mixture about 5minutes.
- 3- Calculate the specific heat of the calorimeter.
- 4- Repeat these steps for other calorimeters.

**Calculation:**

$$Q_{ABS} = -Q_{REL}$$

$$Q_{hot\ water} = -(Q_{cold\ water} + Q_{calorimeter})$$

$$Q_{calorimeter} = -Q_{cold\ water} - Q_{hot\ water}$$

$$(C \cdot \Delta T)_{calorimeter} = -(m \cdot S \cdot \Delta T)_{cold\ water} - (m \cdot S \cdot \Delta T)_{hot\ water}$$

$$\therefore C_{calorimeter} = \frac{-(m \cdot S \cdot \Delta T)_{cold\ water} - (m \cdot S \cdot \Delta T)_{hot\ water}}{\Delta T_{calorimeter}}$$

Note:

$$\text{Specific heat capacity of calorimeter} = \frac{C_{calorimeter}}{\text{mass of calorimeter}}$$

$$C_{cal} = m_{cal} \times S_{cal}$$

**How to Calculate a Calorimeter Constant?**

**Example 1:** When 40.0 mL of water at 60.0 °C is added to 40.0 mL at 25.0 °C water already in a calorimeter, the temperature rises 15.0 °C. What is the calorimeter constant?

**Solution:**

1) Hot water lost:

$$Q = mS\Delta T = 40g \times 4.18 \frac{J}{g \cdot ^\circ C} \times 20^\circ C = 3344$$

2) Cold water got:

$$Q = mS\Delta T = 40g \times 4.18 \frac{J}{g \cdot ^\circ C} \times 15^\circ C = 2508J$$

3) The calorimeter got the rest:

$$3344 - 2508 = 836J$$

4) Find the heat capacity of the calorimeter:

$$C_{cal} = \frac{836J}{15^\circ C} = 55.73 \frac{J}{^\circ C}$$

## **Experiment (6): Calculation of the specific heat of the substance.**

The specific heat of the substance is the number of calories required to raise the temperature of one gram of the substance one degree centigrade.

The experimental determination of the specific heat of a metal by the method of mixtures consists in dropping a known mass of the metal at a known high temperature into a known mass of water at a known low temperature. The equilibrium temperature is then measured. The heat absorbed by the water, calorimeter, and stirrer is equal to the heat lost by the metal.

### **Apparatus and chemicals:**

Cup calorimeter, different metal, oven.

### **Procedure:**

- 1- Heat a metal (Known weight) to 100°C in oven.
- 2- Put the metal in a cup calorimeter which contained 100ml D.W.
- 3- Record the temperature before and after adding the metal

### **Calculation:**

$$Q_{metal} = -(Q_{water} + Q_{cal})$$

**Example:** A piece of metal with a mass of 7.5g is heated to 200°C before being dropped into 100g of water. This causes water temperature to rise from 21.1°C to 30.8°C. What is the metals specific heat capacity?

**Solution:**

$$Q_{metal} = m \cdot S \cdot \Delta T$$

$$\therefore S_{metal} = \frac{Q_{metal}}{m_{metal} \times \Delta T}$$

$$-Q_{metal} = Q_{water}$$

$$Q_{water} = 100 \times 4.18 \times (30.8 - 21.1) = 4054.6 J$$

$$\therefore S_{metal} = \frac{Q_{metal}}{m_{metal} \times \Delta T} = \frac{-4054.6}{7.5 \times (30.8 - 200)}$$

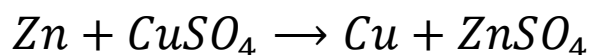
$$S_{metal} = 3.195 \frac{J}{g \cdot K}$$



## **Experiment (7): Determining the enthalpy of reaction of zinc displacing copper from copper (II) sulfate solution**

Enthalpy is the amount of heat that is lost or gained by the atoms in a chemical reaction. Some reactions are endothermic and absorb heat from their surroundings. Some reactions are exothermic and lose heat to their surroundings.

When a more electropositive metal displaces a less electropositive metal from a solution of its salt, heat change occurs.



The literature value for this displacement reaction can be considered as  $-217\text{kJ mol}^{-1}$ .

### **Apparatus and chemicals:**

Calorimeter, Zn powder,  $\text{CuSO}_4$ .

### **Procedure:**

1- Pipette 50.0 ml 0.1 M copper (II) sulphate solution into a 100 ml expanded polystyrene cup. Record the temperature of the solution using a 0-50 °C thermometer.

2- Weigh 0.5 g zinc dust using a clean paper. Add the zinc dust to the copper(II) sulphate solution, at a time. Then, start the stop clock immediately.

3- Stir regularly with the thermometer. Record the temperature every minute for ten minutes.

**Calculation:**

$$Q_{reaction} = -(m \times S \times \Delta T)$$

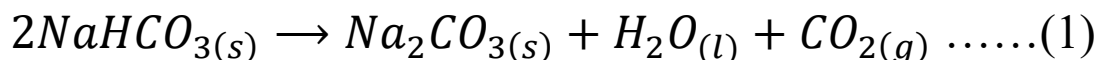
$$Q_{reaction} = -(m \times 4.18 \times \Delta T)$$

$$\Delta H_r = \frac{Q_{reaction}}{n}$$

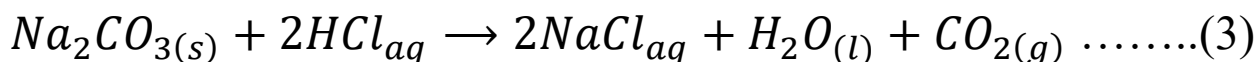
*n = minimum mole in the reaction*

### Experiment (8): Calculation the enthalpy of the decomposition of sodium hydrogen carbonates Using Hess's Law.

The enthalpy change of reaction for the decomposition of sodium hydrogen carbonate,  $NaHCO_{3(s)}$ , cannot be measured directly, then we must carry out two separate experiments and use the results of these experiments to determine the enthalpy change of reaction for the decomposition of sodium hydrogen carbonate.



The two other reactions involve mixing sodium hydrogen carbonate and sodium carbonate with hydrochloric acid.



$$\Delta H_{eq1} = 2 \times \Delta H_{eq2} - \Delta H_{eq3}$$

#### Apparatus and chemicals:

Calorimeter,  $NaHCO_{3(s)}$ ,  $HCl$ ,  $Na_2CO_{3(s)}$ .

**Procedure:**

- 1- Add 2g of  $Na_2CO_{3(s)}$  to 50ml of 0.5M HCl, and then calculate  $\Delta T$ .
- 2- Add 2g of  $NaHCO_{3(s)}$  to 50ml of 0.5M HCl, and then calculate  $\Delta T$ .
- 3- Calculate Q and then  $\Delta H$  of each reaction.
- 4- Calculate  $\Delta H$  of the decomposition of sodium hydrogen carbonates.